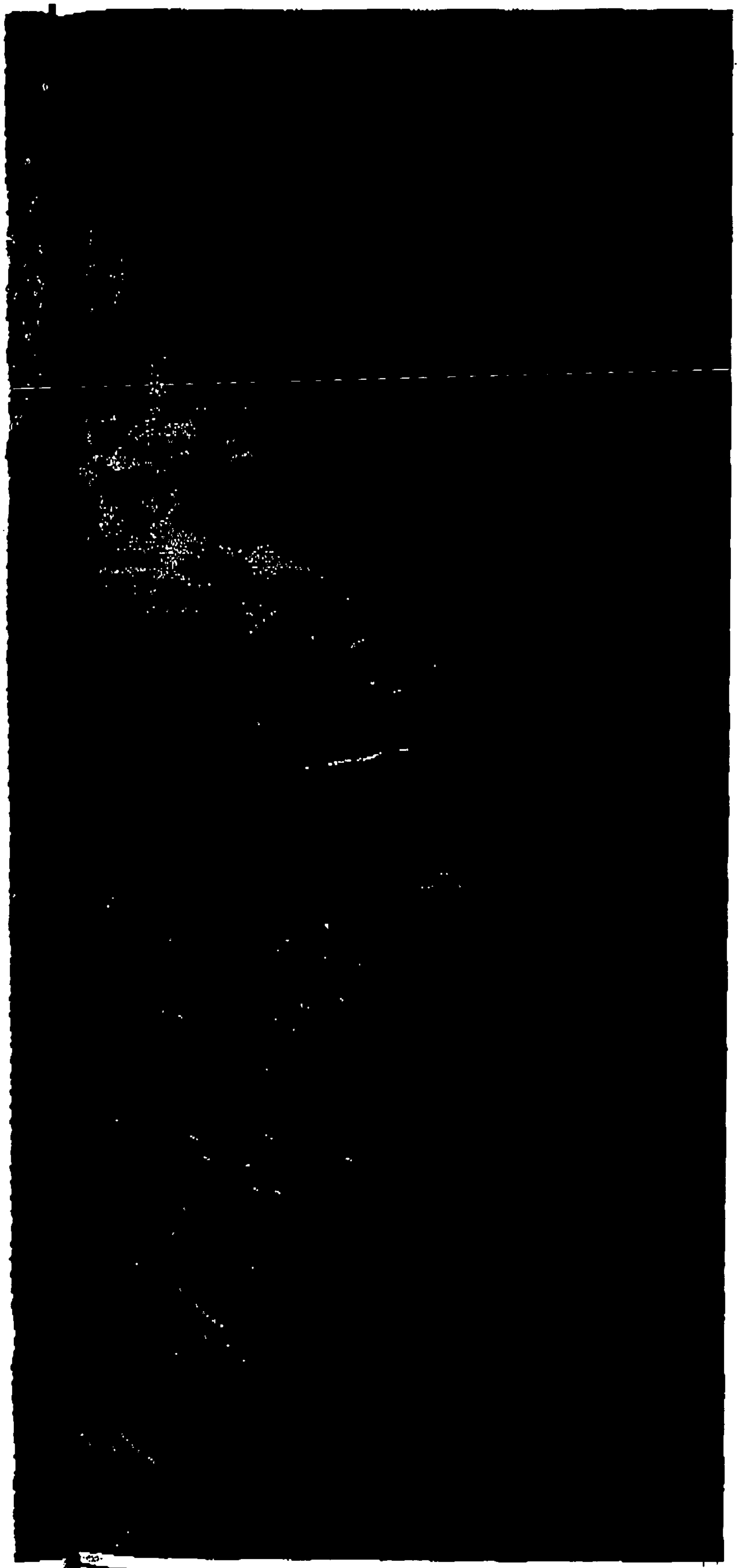
EXHIBIT 2



Chemistry for Engineers and Scientists

LEONARD W. FINE
Columbia University

HERBERT BEALL
Worcester Polytechnic Institute

SAUNDERS COLLEGE PUBLISHING
Philadelphia Chicago Fort Worth
San Francisco Montreal Toronto
London Sydney Tokyo

A LEXAN polycarbonate shield can stop a 357 magnum bullet. Polarized light shows stresses along the fracture lines.

Copyright © 1990 by Saunders College Publishing, a division of Holt, Rinehart and Winston, Inc.

All rights reserved. No part of this publication may be reproduced or transmitted in any form or by any means, electronic or mechanical, including photocopy, recording, or any information storage and retrieval system, without permission in writing from the publisher.

Requests for permission to make copies of any part of the work should be mailed to: Copyrights and Permissions Department, Holt, Rinehart and Winston, Inc., Orlando, FL 32887

Text typeface: Baskerville
Compositor: General Graphic Services
Acquisitions Editor: John Vondeling
Developmental Editor: Sandi Kiselica
Managing Editor: Carol Field
Project Editor: Marc Sherman
Copy Editor: Jay Freedman
Manager of Art and Design: Carol Bleistine
Art and Design Coordinator: Doris Bruey
Text Designer and Layout Artist: Ed Butler
Cover Designer: Lawrence Didona
Text Artwork: J & R Art Services, Inc.
Director of EDP: Tim Frelick
Production Manager: Bob Butler

Cover credit: © Andrew Davidhazy, Rochester Institute of Technology

Printed in the United States of America

CHEMISTRY FOR ENGINEERS AND SCIENTISTS

ISBN 0-03-021537-4

Library of Congress Catalog Card Number: 89-043042

0123 032 987654321

(NH_4Cl) is added to sufficient water to make up one liter of solution.
Answer: $[H_3O^+] = 1.3 \times 10^{-5}$; pH = 4.89; $[OH^-] = 7.7 \times 10^{-10}$. ■

Buffers

When an acid or a base, even in relatively small amounts, is added to water, there is a large change in the pH of the solution. For example, addition of as litle as 0.010 mol of HCl lowers the pH of solvent water from 7 to 2, a change of 5 pH units. A **buffer** is a solution that is capable of absorbing excess acid or base without major changes in pH. Because such a solution must be able to react with a large fraction of added acid or base, the buffer solution must contain both an acid and a base. However, neither can be strong, since they would simply neutralize each other:

$$H_3O^+ \text{ (aq)} + OH^- \text{ (aq)} \longrightarrow 2\ H_2O \text{ (liq)}$$

Instead, a buffer must contain a weak acid and a weak base. More specifically, a buffer must contain a weak acid and its conjugate base.

For example, a liter of solution containing 0.50 mol of formic acid (HCOOH) and 0.50 mol of sodium formate (HCOONa) can provide effective buffering action. Before any dissociation occurs, the resulting solution is 0.50 M in formic acid and sodium formate. To determine the $[H_3O^+]$ and pH of this solution, keep in mind that sodium formate is a freely soluble salt, which dissociates essentially completely:

$$HCOONa \text{ (s)} \longrightarrow Na^+ \text{ (aq)} + HCOO^- \text{ (aq)}$$

Thus, the solution is 0.50 M in formate ion. HCOOH ionizes according to the following equation:

$$HCOOH \text{ (aq)} + H_2O \text{ (aq)} \rightleftharpoons H_3O^+ \text{ (aq)} + HCOO^- \text{ (aq)}$$

$$K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]} = 1.8 \times 10^{-4}$$

$$[H_3O^+] = (1.8 \times 10^{-4}) \frac{[HCOOH]}{[HCOO^-]}$$

If we let $[H_3O^+]$ formed as a result of this ionization be x, then, assuming $\ll 0.500$

$$x = [H_3O^+] = (1.8 \times 10^{-4}) \left(\frac{0.500}{0.500}\right) = 1.8 \times 10^{-4} \text{ M}$$

$$\text{pH} = -\log [H_3O^+] = -\log (1.8 \times 10^{-4}) = 3.74$$

Now suppose we dilute this buffer solution by adding enough water o make 2.0 L. The concentrations of HCOOH and $HCOO^-$ drop accordngly, to 0.25 M, and we may write the equilibrium expression as follows:

$$[H_3O^+] = (1.8 \times 10^{-4}) \left(\frac{0.25}{0.25}\right) = 1.8 \times 10^{-4} \text{ M}$$

Note that the $[H_3O^+]$ and pH are unchanged, demonstrating one very important feature of buffer solutions: moderate dilution of a buffer sotion does not significantly change its pH.

Let's now go back to our original, undiluted 1.0 L of buffer solution nd add 0.010 mol of strong acid. What effect does this have on the $[H_3O^+]$ nd pH of the resulting solution? Remember what happened when we